TOPIC 4
MATTER

4.1 States of matter

4.1.1 Arrangements of particles

All matter is made of up tiny particles (electrons, ions, atoms, molecules) that are in perpetual random motion. The large-scale properties of matter, including density, pressure, temperature, state (and others such as 'hardness' or 'runniness') are related to the way these particles are arranged and the way they move.

By the **state of matter** (or **phase of matter**) we mean whether it is a solid, liquid or gas. Any substance can exist in *any* of these states, depending on its temperature and pressure. For example, oxygen and carbon dioxide are gases at normal room temperature and pressure, but if oxygen is cooled to –183 °C it liquefies and if carbon dioxide is cooled by a similar amount it solidifies (to become so-called dry ice). Iron is solid at room temperature and pressure but if heated to 1535 °C it melts to a liquid and at 3027 °C it becomes a gas.

In the **solid** state, the atoms or molecules are packed close together in fixed positions and their only motion is vibration (see Figure 4.1, the spoon). The fixed arrangement of particles means that a solid keeps its shape. In **metals** and some other solids known as semiconductors, electrons are able to move freely throughout the material, enabling them to conduct electricity. If the particles are arranged in regular rows and layers, the solid is then a **crystal**; metals and minerals have a crystalline structure. A single crystal has a regular shape that depends on the arrangement of the particles.

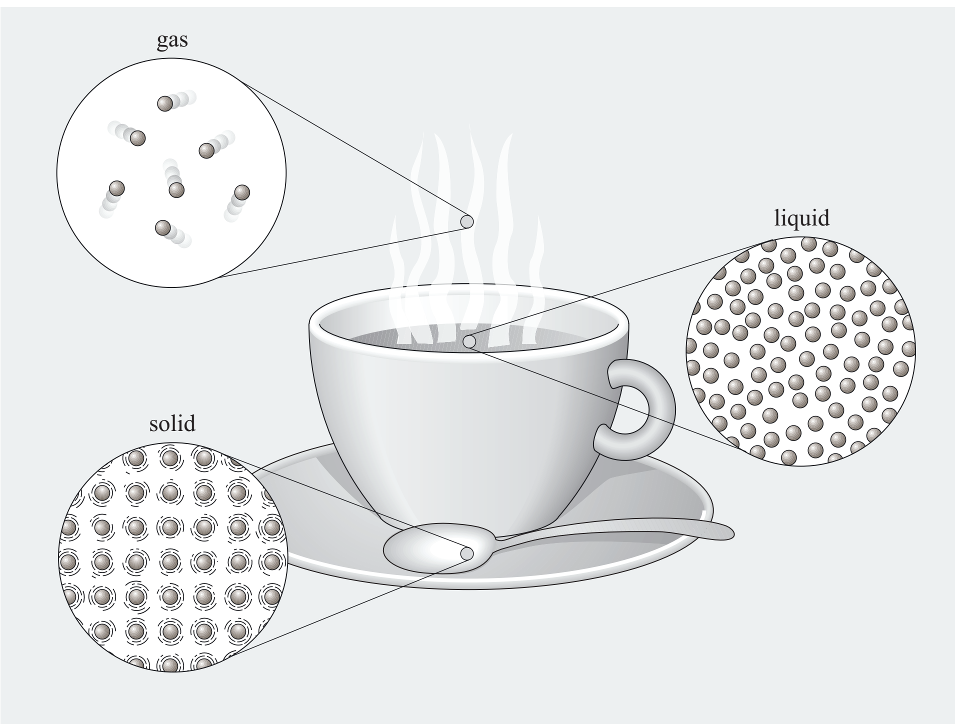


Figure 4.1 Schematic arrangements of particles in a solid (the spoon), a liquid (the tea), and a gas (the 'steam', or vapour).

The particles in **liquid** are also close together — about as close as those in a solid — only now they are not fixed in place but can move around freely (Figure 4.1, the tea) as they have more energy than the particles of a solid. Liquids therefore flow to take up the shape of their container (in this case, a cup) while maintaining a constant volume.

A **gas** is characterized by being made up of atoms or molecules that are widely separated and moving around at high speed (Figure 4.1, the steam), colliding frequently with one another. A gas will therefore flow and expand its volume rapidly to fill any container in which it is placed. Since both liquids and gases can flow freely, they are collectively known as **fluids**.

Finally we come to **plasma**, sometimes called the *fourth* state of matter. A plasma is a gas composed largely of ions and electrons which are free to move independently. As these particles have electric charge, they interact with one another via electrostatic forces, and this considerably affects the properties of the substance. Stars, and many regions of the interstellar medium, consist largely of plasma.

QUESTION 4.1

What is wrong with each of the following statements?

(i) Aluminium is a solid.

(ii) Nitrogen is a gas.

(iii) All gases are plasmas.

4.1.2 Density

Solids, liquids and gases are, loosely speaking, characterized by having different **mass density**, (generally just called the **density**) and this in turn is related to the way their particles are arranged. The density of a substance is its mass per unit volume; in SI units, it is the mass of 1 m^3 of the substance, so the SI units of density are $kg\,m^{-3}$. Given the mass m and volume V of a sample, its density, usually denoted by the Greek letter rho, ρ, can be calculated from

$$\rho = \frac{m}{V} \qquad (4.1)$$

The densities of the liquid phase and the solid phase of a given substance are similar. On changing between the solid and liquid states the separation of the particles changes very little, so the same mass occupies a very similar volume.

Gases, in general, have much lower densities than solids or liquids, as their particles are much further apart and so a given mass of gas occupies a much larger space than the same mass of that substance when it is in the solid or liquid state. For example, liquid water at its boiling temperature and normal atmospheric pressure has a density of $1 \times 10^3\,kg\,m^{-3}$, whereas steam at the same temperature and pressure has a density of only about $0.6\,kg\,m^{-3}$.

EXAMPLE 4.1

If the average density of the materials that make up the planet Jupiter is $1.33 \times 10^3\,\mathrm{kg\,m^{-3}}$, and Jupiter's volume is $1.43 \times 10^{24}\,\mathrm{m^3}$, what is Jupiter's mass?

From Equation 4.1

$$m = \rho V$$
$$= 1.33 \times 10^3\,\mathrm{kg\,m^{-3}} \times 1.43 \times 10^{24}\,\mathrm{m^3} = 1.90 \times 10^{27}\,\mathrm{kg}$$

QUESTION 4.2

A certain meteorite has a mass of 2.34 kg and a volume $5.22 \times 10^{-4}\,\mathrm{m^3}$. What is its density?

When we are dealing with the behaviour of individual particles, particularly when densities are very low (for example in the interstellar medium), it is often useful to characterize substances by their **number density**, i.e. the *number* of particles per unit volume. Number density is usually denoted n, and has SI units $\mathrm{m^{-3}}$. A word of warning: sometimes we use number density to mean the total number of particles per unit volume (regardless of their nature) whereas at other times we are interested in the number density of a particular substance, e.g. the number of hydrogen atoms per unit volume — this is usually clear from the context.

Number density and mass density are closely related. If there are n particles per unit volume, and each particle has a mass M, then the overall mass per unit volume is nM, in other words

$$\rho = nM \tag{4.2}$$

EXAMPLE 4.2

In a certain region of the interstellar medium, the number density of hydrogen atoms is $1.00 \times 10^{11}\,\mathrm{m^{-3}}$. Given that the mass of a hydrogen atom is $1.67 \times 10^{-27}\,\mathrm{kg}$, what is the mass density of hydrogen in this region?

From Equation 4.2, $\rho = 1.00 \times 10^{11}\,\mathrm{m^{-3}} \times 1.67 \times 10^{-27}\,\mathrm{kg} = 1.67 \times 10^{-16}\,\mathrm{kg\,m^{-3}}$.

QUESTION 4.3

In the Sun's core, the mass density is about $1 \times 10^5\,\mathrm{kg\,m^{-3}}$. If you assume that the Sun's core is mainly composed of *protons*, each with mass $1.67 \times 10^{-27}\,\mathrm{kg}$, what is the number density of protons in the Sun's core?

4.2 Temperature and pressure

4.2.1 Kinetic energy and absolute temperature

The average kinetic energy of an object's constituent particles is related to its **temperature**. As temperature rises, the average particle kinetic energy increases. Likewise, as an object cools, so the average kinetic energy of its particles decreases. Eventually a point is reached at which the particles can lose no more energy and no further cooling can be achieved. This point is the same for all substances, and the temperature at which this occurs is known as **absolute zero**

— the lowest temperature possible. On the Celsius temperature scale (formerly known as centigrade) this temperature has a value of −273.15 °C.

For scientific purposes it is useful to define a temperature scale that starts at absolute zero. This is the **absolute temperature scale**, also known as the **Kelvin scale** after physicist and engineer William Thomson, Lord Kelvin (1824–1907). The unit of temperature on this scale is the kelvin (K) (not 'degree kelvin', and not °K), and the kelvin is the SI unit of temperature. Figure 4.2 shows the relationship between the Kelvin and Celsius temperature scales. A kelvin is the same size as a °C, so to convert from °C into K you just add 273.15 to the Celsius temperature:

$$(\text{temperature in K}) = (\text{temperature in °C}) + 273.15 \qquad (4.3\text{a})$$

or, in symbols

$$T/\text{K} = T/°\text{C} + 273.15 \qquad (4.3\text{b})$$

There is a direct relationship between the particles' kinetic energy and the absolute temperature. The simplest case is that of a gas made up of single atoms, where the only contribution to the internal energy is the motion of each atom as a whole. The atoms will transfer energy to one another as they collide, so it is not possible to pinpoint the kinetic energy of any particular one, but the total amount of energy that they share between them depends only on the temperature. At an absolute temperature T, the average kinetic energy of each particle, E_k is given by

$$E_k = \frac{3kT}{2} \qquad (4.4)$$

where k is the **Boltzmann constant**, 1.38×10^{-23} J K^{-1}, named after Austrian physicist Ludwig Boltzmann (1844–1906). If a gas is made up of two or more types of atom, then the average kinetic energy per atom is still given by Equation 4.4, but the less massive atoms will on average be moving faster in order to have the same kinetic energy (E_k also equals $mv^2/2$ so if m is small then v must be large to give the same E_k).

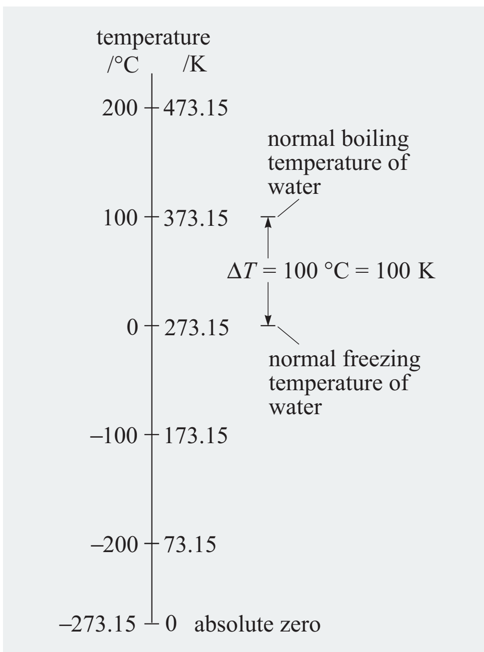


Figure 4.2 The Celsius and Kelvin temperature scales.

In gases with molecules that can rotate and vibrate as well as moving to and fro, there are other contributions to the internal energy but Equation 4.4 still describes the so-called translational kinetic energy, i.e. the kinetic energy of the to-and-fro motion.

For other types of substance (solids for example, or plasmas) the relationship between temperature and internal energy is not quite the same as that shown in Equation 4.4, but there is still a direct connection between temperature and the energy of the particles.

EXAMPLE 4.3

What is the average kinetic energy of an atom in a gas whose temperature is 50 °C?

First express T in SI units of kelvin: from Equation 4.3, $T = 323.15\,\mathrm{K}$

Then from Equation 4.4,

$$E_\mathrm{k} = \frac{3\times 1.38\times 10^{-23}\ \mathrm{J\,K^{-1}}\times 323.15\,\mathrm{K}}{2} = 6.7\times 10^{-21}\ \mathrm{J}$$

QUESTION 4.4

The temperature near the top of the Earth's atmosphere is about 1000 K. What is this temperature in °C? What is the average kinetic energy of atoms in the atmosphere here?

QUESTION 4.5

The mass of a nitrogen atom is 14 times that of a hydrogen atom, and that of an oxygen atom is 16 times that of a hydrogen atom. In a gas that is a mixture of hydrogen, oxygen and nitrogen atoms, which atoms will have, on average, the most kinetic energy? Which atoms will, on average, have the highest speeds?

4.2.2 Pressure

Pressure, p, is defined as force per unit area

$$p = \frac{F}{A} \tag{4.5}$$

where A is the area (measured 'square on') over which a force F is acting. The SI unit of pressure is the pascal, Pa, named after the French scientist Blaise Pascal (1623–1662). Note that $1\,\mathrm{Pa} = 1\,\mathrm{N\,m^{-2}}$.

EXAMPLE 4.4

If your weight is 600 N and the area of your shoe soles in contact with the floor is $0.020\,\mathrm{m^2}$, what is the pressure you exert on the floor? If you lift up one foot so that you halve the area in contact with the floor, what happens to the pressure you exert?

$$p = \frac{600\,\mathrm{N}}{0.020\,\mathrm{m^2}} = 3.00\times 10^4\ \mathrm{Pa}$$

If you halve the area the pressure is doubled to $6.00 \times 10^4\,\mathrm{Pa}$.

The pressure of the atmosphere, which is greatest at sea-level and decreases with height, and the increase of pressure with depth under water, can be explained in terms of the weight of the overlying air and water. However, within a fluid, the random motion of the particles ensures that pressure is exerted equally in all directions and is not associated with any particular direction.

4.2.3 Gas pressure

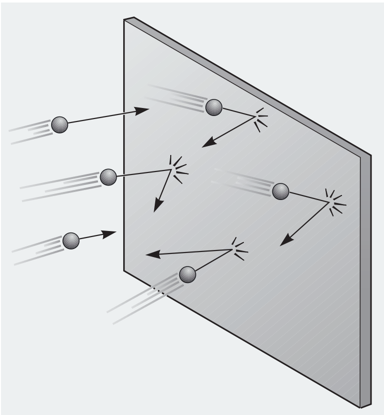


Figure 4.3 A gas exerts a force on a surface as a result of collisions.

The continuous motion of gas molecules gives rise to a pressure as they bombard any surfaces that they strike, as shown schematically in Figure 4.3.

The pressure, i.e. the force on a given area of surface, can be increased in two ways. First, if the number density, n, of gas particles increases then there are more frequent collisions. Second, if the temperature, T, increases then the particles move faster and the collisions not only become more frequent but each collision is more violent so giving rise to a greater force. These relationships can be summarized:

$$p = nkT \tag{4.6}$$

where k is the Boltzmann constant, $1.38 \times 10^{-23}\,\mathrm{J\,K^{-1}}$ and the temperature, T, is the absolute temperature.

EXAMPLE 4.5

The pressure near the top of the Earth's atmosphere is much less than that at sea-level, and yet the temperature is much greater. What can you deduce about the number density of gas particles near the top of the atmosphere?

The number density must be much less than at sea-level. If it were the same, then the higher temperature would give rise to a higher pressure.

EXAMPLE 4.6

At sea-level, the pressure of the Earth's atmosphere is 1.0×10^5 Pa and the number density of gas molecules is about $2.5 \times 10^{25}\,\mathrm{m^{-3}}$. What do these values predict for the temperature at the Earth's surface?

Rearranging Equation 4.6 to make T the subject

$$T = \frac{p}{nk} = \frac{1.0\times10^{5}\ \mathrm{Pa}}{(2.5\times10^{25}\ \mathrm{m^{-3}})\times(1.38\times10^{-23}\ \mathrm{J\,K^{-1}})} = 290\ \mathrm{K}\ (=17\,°\mathrm{C})$$

QUESTION 4.6

At the surface of Venus the particle number density in the atmosphere is about $9.0 \times 10^{26}\,\mathrm{m^{-3}}$ and the temperature is about 733 K. What would you predict for the atmospheric pressure at this location?

4.3 Atoms and their constituents

4.3.1 Atomic structure

Atoms are the basic components of matter. They come in many types but all are very small, about 2×10^{-10} m in diameter. An atom consists of a cloud of **electrons**, each with a negative electric charge, surrounding a tiny positively charged **nucleus**. The nucleus has a diameter of only about 1×10^{-14} m but accounts for nearly all the mass of an atom. In general a nucleus consists of a tightly-bound mixture of **neutrons** and **protons**. The structure of an atom is shown schematically in Figure 4.4. While in some ways, this picture is quite unrealistic (the electrons are not confined to narrow orbits, for example, but rather form a fuzzy cloud of charge around the nucleus) it does represent many key features of atomic structure.

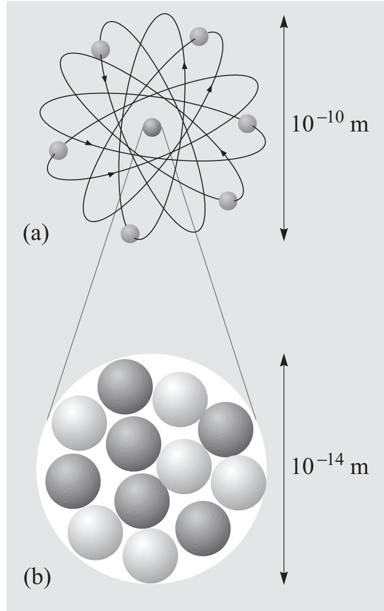


Figure 4.4 A schematic representation of a typical atom: (a) electrons move around the nucleus, (b) the nucleus is made up of protons and neutrons.

The masses and electric charges of the electron, proton and neutron are given in Table 4.1. Notice that the charges of the electron and proton are *exactly* the same size but of opposite sign. An atom contains equal numbers of protons and electrons, so is electrically neutral. However, an atom can gain or lose electrons in which case it is called an **ion**. The process of losing electrons is called **ionization**.

Table 4.1 The particles that make up an atom.

Particle and symbol	mass/kg	charge/C
electron e or e^-	9.109×10^{-31}	-1.602×10^{-19}
proton, p	1.6726×10^{-27}	$+1.602 \times 10^{-19}$
neutron, n	1.6749×10^{-27}	zero

There are many types of atom, each being characterized by the number of protons and neutrons in its nucleus. All the atoms of a particular **chemical element**, represented by its own unique chemical symbol, have the same number of protons in the nucleus. The number of protons is called the **atomic number** and given the symbol Z. The number of protons in turn determines the number of electrons in the atom, and it is the electrons that take part in chemical reactions. It is the electrons that give an atom its particular chemical properties.

Ions are denoted by a chemical symbol with a superscript indicating the net change in terms of the proton's charge. So if an atom of aluminium loses an electron it becomes Al^+, if it loses two electrons it becomes Al^{2+}, and if it gains an electron it is Al^-.

QUESTION 4.7

(i) Write down the chemical symbol for the ion created when an atom of iron loses two electrons. (ii) What must happen to an atom of chlorine, Cl, in order to create the ion Cl^-?

4.3.2 Isotopes

Neutrons and protons are collectively called **nucleons**. The total number of nucleons in an atom is called the **mass number** and given the symbol A. To a good approximation, a proton and a neutron have the same mass, 1.67×10^{-27} kg, and so the mass of a nucleus is very close to A times this mass. As the electron's mass is so very much smaller than that of a nucleon the mass of the nucleus is, for most purposes, the same as the mass of the atom.

Nuclei of any given element, characterized by its **atomic number** Z, can have various numbers of neutrons and hence various different values of mass number A. These different nuclei are all **isotopes** of the same element. Different isotopes of an element will all undergo the same chemical reactions but will have different nuclear reactions.

A particular isotope of a particular element corresponds to a particular type of nucleus, called a **nuclide**, and is specified by its values of A and Z. The usual way of representing a nuclide is ${}^{A}_{Z}\mathrm{X}$, where X is the chemical symbol. For example, one isotope of aluminium is ${}^{27}_{13}\mathrm{Al}$. All isotopes of aluminium have 13 protons, and this particular one also has 14 neutrons giving it a mass number of 27. The isotope is often referred to as aluminium-27 in order to distinguish it from another isotope aluminium-26 which has only 13 neutrons and hence a mass number 26. Quite often we omit the atomic number from the symbol, ${}^{27}\mathrm{Al}$, because it adds no information that is not already implied by the chemical symbol, although when writing nuclear reaction equations it is useful to include it.

The most common isotope of the lightest element, hydrogen, has a nucleus consisting of just a single proton. There are two less-common hydrogen isotopes: **deuterium**, also called 'heavy hydrogen', has a neutron and a proton in the nucleus, and **tritium**, has a proton and two neutrons. These isotopes are unusual in that they are sometimes given their own chemical symbols. Deuterium can be represented as either ${}^{2}_{1}\mathrm{H}$ or ${}^{2}_{1}\mathrm{D}$, while tritium can be either ${}^{3}_{1}\mathrm{H}$ or ${}^{3}_{1}\mathrm{T}$.

QUESTION 4.8

An isotope of iron is represented by ${}^{56}_{26}\mathrm{Fe}$. How many protons are in its nucleus? How many neutrons?

QUESTION 4.9

Barium (Ba) has $Z = 56$. Write down the symbol for the isotope of barium that contains 81 neutrons.

4.3.3 Atomic energy levels

An electron and an atomic nucleus have kinetic energy and potential energy. The potential energy depends on the distance between the electron and the nucleus and arises because a nucleus and an electron have opposite electric charge and so attract one another by the electric force. Just as increasing the separation between an object and the Earth increases their gravitational potential energy, so increasing the separation of particles interacting by the electric force also increases their electric potential energy. The total internal energy of an atom is the sum of the kinetic and potential energies due to the motion and positions of the nucleus and all the electrons.

Within any atom, there can only be certain discrete energies. These permitted **energy levels** are unique for each chemical element and are determined according to the laws of quantum physics that become important when we are dealing with matter on very small scales. The simplest case is that of hydrogen, which has only one electron. Its energy levels are shown in Figure 4.5. They are labelled $n = 1, 2, 3$ … starting with the lowest level, known as the **ground state**. This corresponds to the electron being close to the nucleus. The higher levels correspond to the atom being in an **excited state**, meaning that it has energy over and above its minimum value.

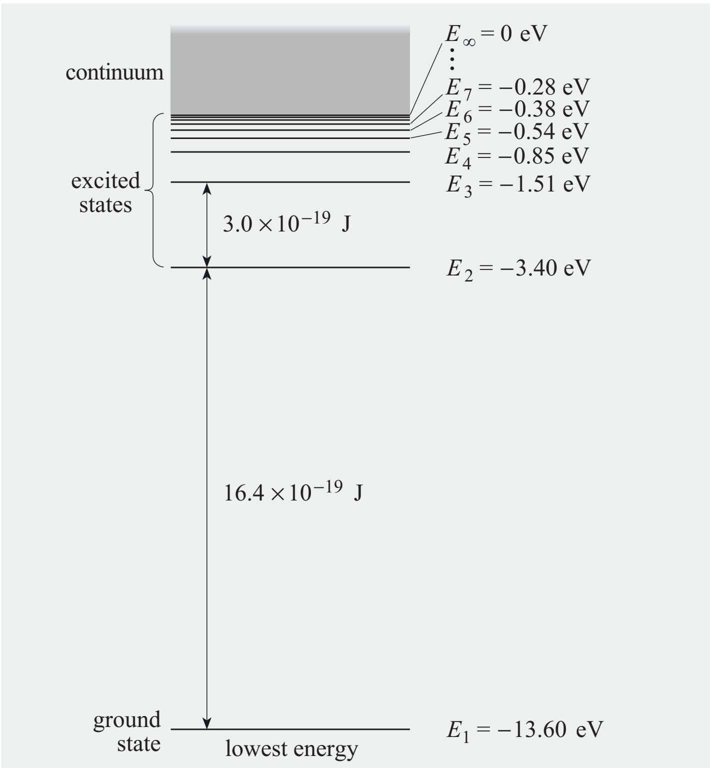


Figure 4.5 Energy levels of a hydrogen atom.

Notice that the energy levels in Figure 4.5 can be labelled either in SI units of joules or in electronvolts, eV. Notice, too, that the energies are all negative. This arises because it is convenient to define the energy to be zero when the electron is just detached from the nucleus. When it is attracted towards the nucleus it loses energy (like an apple falling to Earth loses gravitational potential energy) so all the energy levels within the atom must be negative.

An atom can only change its energy if it gains or loses exactly the right amount to make a transition to another energy level. The most common way to make a transition from an excited state to one of lower energy is for the atom to emit a photon. The photon energy E_{ph} is given by

$$E_{ph} = E_{high} - E_{low} \quad (4.7)$$

where E_{high} is the energy of the higher state and E_{low} that of the lower.

EXAMPLE 4.7

If a hydrogen atom makes a transition from level $n = 2$ to $n = 1$, what is the energy of the photon emitted?

The energy of the higher level is $E_{high} = -3.40$ eV and that of the lower level is $E_{low} = -13.6$ eV.

From Equation 4.7

$$E_{ph} = -3.40\,\text{eV} - (-13.60\,\text{eV}) = 10.2\,\text{eV}$$

Since photon energy is related to the frequency and wavelength of electromagnetic radiation, each element has its own **emission spectrum**, in other words its own unique set of wavelengths that its excited atoms can emit. Most transitions between atomic energy levels involve radiation in the visible, infrared or ultraviolet region of the electromagnetic spectrum.

If an atom absorbs a photon whose energy exactly corresponds to the difference between its current energy level and a higher one, then **excitation** occurs, i.e. it makes a transition to the higher level. If a beam of radiation with a continuous range of wavelengths shines through a sample of atoms in low energy states, then photons will be absorbed if their energy corresponds exactly to the differences between energy levels. Radiation of wavelengths corresponding to these energies will therefore be removed from the beam, and an **absorption spectrum** will be observed. That is, there will be much less radiation at those wavelengths. These missing wavelengths are exactly the same as the wavelengths of the emission spectrum.

A more common way for atoms to become excited is by **collisional excitation**, i.e. in collisions with other atoms or with electrons. If the kinetic energy of the colliding particles is enough to cause excitation then some of their energy may be transferred to the atom and they will move apart with reduced kinetic energy.

If the atom receives sufficient energy, it is possible to dislodge an electron completely and produce an ion. The minimum energy needed to produce an ion, starting with the atom in its ground state, is called the ionization energy. On Figure 4.5, this corresponds to a transition from the level $n = 1$ to the top energy level $n = \infty$ or $E = 0$. If any more energy is transferred to the atom, then the transition corresponds to ending up in the region labelled 'continuum' on Figure 4.5, which means the electron and ion have some kinetic energy that enables them to move apart independently of one another.

EXAMPLE 4.8

In a certain region of the interstellar medium, the average kinetic energy of the particles is 12 eV. Are hydrogen atoms likely to become excited in collisions?

Yes they are. Excitation to the $n = 2$ level requires 10.2 eV (see Example 4.7) so collisions between particles with kinetic energy 12 eV are sufficiently energetic to cause excitation.

EXAMPLE 4.9

From Figure 4.5, what is the ionization energy of hydrogen?

To reach the energy level with $E = 0$, starting from the ground state, the atom must receive 13.60 eV or 21.79×10^{-19} J.

QUESTION 4.10

A hydrogen atom makes a transition from the energy level $n = 2$ to the level $n = 3$. State whether this will involve the emission or the absorption of a photon, and calculate its energy in eV.

QUESTION 4.11

At room temperature, the average particle kinetic energy is about 4×10^{-2} eV, or about 6×10^{-21} J (see Example 4.3). Explain whether hydrogen atoms at room temperature are likely to be in an excited state.

QUESTION 4.12

Explain what will happen if a beam of photons, all with energy 11.0 eV, shines onto a sample of hydrogen atoms (a) if all the atoms are initially in the $n = 2$ energy level (b) if all the atoms are initially in the ground state.

4.3.4 Subatomic particles

Atoms were at first thought to be **fundamental particles** — that is, the simplest building blocks from which all matter is constructed. It was thought that atoms were indivisible and did not contain any other particles. Then in the late 19th and early 20th centuries a different picture emerged: atoms are complex particles made up of protons, neutrons and electrons. The particles that make up atoms are **subatomic particles** — particles that are simpler than atoms.

Research during the 20th century revealed the existence of other subatomic particles, which are not found as constituents of atoms but are produced in radioactive decay, and/or detected in cosmic radiation (high-energy particles that enter the Earth's atmosphere and have extraterrestrial origin) and/or created in specially built particle accelerators, where mass–energy equivalence leads to the production of particles in high-energy collisions. Such particles would also be present in the very energetic conditions of the very early universe.

But are subatomic particles fundamental or are they in turn made up of simpler particles?

Electrons *are* believed to be fundamental. They belong to a family of six fundamental particles called **leptons**, listed in Table 4.2. The muon (whose symbol is the Greek letter mu, μ) and tauon (Greek tau, τ) are similar to the electron but have greater mass. The muon has about 200 times the mass of the electron and the tauon's mass is about 3500 times that of the electron. The superscript minus signs indicate that these particles have negative electric charge equal to the electron charge $-e = -1.60 \times 10^{-19}$ C.

Table 4.2 The leptons.

1st generation	2nd generation	3rd generation	Charge
electron, e^-	muon, μ^-	tau, τ^-	$-e$
electron-neutrino, ν_e	muon-neutrino, ν_μ	tau-neutrino, ν_τ	0

The three **neutrinos** (Greek nu, ν) associated with each of the electron, muon and tauon, are uncharged. The neutrinos have *extremely* small masses. It is currently (2002) believed that their masses are not quite zero but they have not yet been reliably determined. The three pairs (e.g. the electron and its neutrino) are often referred to as the three 'generations' of lepton.

To each lepton there is a corresponding **antiparticle**, an **antilepton**. An antiparticle is one with exactly the same mass as the corresponding particle but with charge (and some other properties) exactly opposite. The first antiparticle to be discovered was the antielectron, which is also called the **positron**. Table 4.3 lists the antileptons. The antileptons are denoted by symbols with superscript plus signs (indicating their charge) and the **antineutrinos** have a bar over the nu.

Table 4.3 The antileptons.

1st generation	2nd generation	3rd generation	Charge
positron, e^+	antimuon, μ^+	antitau, τ^+	$+e$
electron-antineutrino $\overline{\nu}_e$	muon-antineutrino, $\overline{\nu}_\mu$	tau-antineutrino, $\overline{\nu}_\tau$	0

The proton and neutron, however, are *not* fundamental particles. Each is made up of three **quarks**, which *are* believed to be fundamental. As with leptons, there are three 'generations' of quarks, which are shown in Table 4.4. The first generation of quarks have the lowest mass and the third are the most massive. Also like leptons, each quark has a corresponding **antiquark**, as listed in Table 4.5. Sometimes the symbols q and $\overline{q}$ are used to represent quarks and antiquarks in general without specifying a particular type.

Table 4.4 The quarks.

1st generation	2nd generation	3rd generation	Charge
up, u	charm, c	top, t	$+2e/3$
down, d	strange, s	bottom, b	$-e/3$

Table 4.5 The antiquarks.

1st generation	2nd generation	3rd generation	Charge
$\overline{u}$	$\overline{c}$	$\overline{t}$	$-2e/3$
$\overline{d}$	$\overline{s}$	$\overline{b}$	$+e/3$

Quarks have never been observed in isolation. They only seem to occur bound together, and particles made up of bound quarks are collectively known as **hadrons**. There are three sorts of hadron: three quarks can join to form a type of particle called a **baryon**, three antiquarks can form an **antibaryon**, and a quark and an antiquark can form a type of particle called a **meson**. All hadrons have charges that are either zero or whole-number multiples of the electron charge. All baryons have antiparticles, which are formed from the corresponding antiquarks.

EXAMPLE 4.10

A proton is composed of two ups and a down quark (uud), giving a total charge of $Q = 2e/3 + 2e/3 - e/3 = +e$, while a neutron has one up and two downs (udd), giving a net charge of zero.

The proton and neutron are both baryons, so the matter that makes up our Earth, and the other matter that we observe in the universe, is sometimes referred to as *baryonic matter*.

Apart for the proton and neutron, there are many other possible hadrons, though none is found as a constituent of ordinary matter. For example, a d quark and a $\overline{\mathrm{u}}$ antiquark form a meson called a π^-, which is found in cosmic radiation but is unstable so decays in a fraction of a second.

QUESTION 4.13

What is wrong with each of the following statements?

(i) Neutrons are fundamental particles.

(ii) Leptons combine to make hadrons.

(iii) All combinations of quarks and/or antiquarks are called baryons.

QUESTION 4.14

What are the constituents of an antiproton? What is the charge of an antiproton?

4.4 Chemical compounds

4.4.1 Symbols and formulae

A chemical element is a substance that is made up of just one type of atom. Each element has its own unique **chemical symbol** consisting of either one or two letters which are generally an abbreviation of its English or Latin name. The first letter of the symbol is always a capital, but the full name of the element is written in lower-case. For example, the symbol for aluminium is Al, while iron is Fe from the Latin *ferrum*. The symbol (e.g. Al or Fe) can stand either for a particular chemical element or for a single atom.

The **Periodic Table** (Table 7.2) lists all the known elements in order of increasing atomic number together with their symbols (and sometimes along with other properties). The layout of the table is such that elements listed in the same column have similar chemical properties.

A **chemical compound** is one in which two or more different types of atom are joined to create a different substance. The properties of a compound are generally quite unlike those of the elements from which it is made. For example, water is a compound of hydrogen and oxygen. The elements hydrogen and oxygen are both *gases* at room temperature while water is a liquid, and we need to breathe oxygen in order to stay alive but breathing water vapour (steam) is no substitute.

The **chemical formula** (or just the **formula**) of a substance is a shorthand way of describing its **chemical composition** — the different atoms it contains and the proportions in which they are present. Chemical symbols represent the atoms the substance contains and subscripts after the symbols indicate their numerical proportions.

EXAMPLE 4.11

The chemical formula for water is H_2O, showing that it is made up of hydrogen (H) and oxygen (O). The subscript 2 after the H means that there are two hydrogen atoms for each atom of oxygen.

EXAMPLE 4.12

Carbon monoxide has the symbol CO while carbon dioxide is CO_2. Both contain carbon (C) and oxygen (O), but the formula CO implies that there are equal numbers of carbon and oxygen atoms while in CO_2 the subscript 2 means that there are two oxygen atoms to each carbon atom.

EXAMPLE 4.13

The mineral haematite is a form of iron oxide and has the chemical formula Fe_2O_3. For every two atoms of iron (Fe) there are three of oxygen (O).

EXAMPLE 4.14

At room temperature, atoms of most gaseous elements tend to join together in pairs to make diatomic molecules. Oxygen gas has the formula O_2 to show that there are two atoms of oxygen joined together.

These examples show that the names of compounds can indicate their make-up (though that is not always the case). For example, compounds in which an element is combined with oxygen are collectively known as **oxides**, and the 'di' in dioxide indicates that there are two oxygens to each atom of the other element. But some oxides, such as water, are better known by other names.

As Example 4.12 indicates, the nature of the compound depends not only on the atoms it contains but also on the proportions in which they are combined — carbon dioxide is a normal part of the air we breathe, but even small quantities of carbon monoxide are toxic.

Example 4.13 also illustrates a convention when writing chemical formulae: if they contain a *metal* that symbol usually comes first.

QUESTION 4.15

The chemical formula for the compound calcium carbonate is $CaCO_3$. What are the numerical proportions of the atoms of calcium (Ca) carbon and oxygen that it contains?

QUESTION 4.16

Write down a chemical formula for the compound called silver perchlorate, which contains equal numbers of atoms of silver (Ag) and chlorine (Cl) and four atoms of oxygen for every atom of silver.

4.4.2 Chemical reactions

A **chemical reaction** is any process in which atoms become joined and/or separated to produce a different substance. Examples of chemical reactions include burning petrol in a car engine to produce exhaust gases, and the rusting or tarnishing of some metals when they are exposed to the mixture of substances that make up the air. Processes such as vaporization or dissolving are *not* chemical reactions because they do not produce different substances: when liquid water boils to become steam, it is still H_2O, and when sugar dissolves in water it is still present as sugar — it can be detected by its taste.

In any chemical reaction, atoms can only ever be rearranged and joined together in different ways; they are never created or destroyed. The atoms contained in the **products** of a reaction (what comes out) must be exactly the same as those contained in the **reactants**, as the starting substances are known.

A chemical reaction can be represented in a **chemical equation** (or reaction equation) using chemical symbols. Since the total number of atoms of each type is unchanged by a reaction, the total number of times each element's symbol appears (taking account of superscripts) must be the same on each side of a reaction equation, that is the equation must be balanced. For example, the most stable form of the element carbon is a black solid at room temperature (soot and charcoal are mainly carbon mixed with a few other substances). When it is heated with carbon dioxide, carbon monoxide is made. The equation for this reaction is

$$C + CO_2 \longrightarrow 2CO \tag{4.8}$$

The left-hand side of Equation 4.8 shows that when a single atom of carbon reacts with carbon dioxide to produce carbon monoxide, a total of two carbon atoms and two oxygen atoms are involved. Before the reaction, one carbon atom is present as pure carbon, while the other is combined with two oxygen atoms. After the reaction, the only substance present is the compound carbon monoxide (CO) and the 2 shows that there are two carbon and two oxygen atoms combined in this way so there are still two atoms of each type present, but they are arranged differently.

To produce a balanced equation from a description of a chemical reaction, it is best to start off by writing down the reactants and products with their correct formulae then adjust the numbers on each side to get a balance.

EXAMPLE 4.15

If the element copper (Cu) is heated in oxygen (O_2) there is a reaction that produces copper oxide (CuO) (which is a black powder). To produce a balanced equation, first write down

$$Cu + O_2 \longrightarrow CuO$$

There are two oxygen atoms on the left but only one on the right, so replace CuO by 2CuO:

$$Cu + O_2 \longrightarrow 2CuO$$

Now the oxygen atoms balance but there are two copper atoms on the right but only one on the left, so replace Cu by 2Cu to get a correctly balanced equation:

$$2Cu + O_2 \longrightarrow 2CuO \tag{4.9}$$

QUESTION 4.17

One of the main constituents of natural gas is the compound methane (CH_4). When methane burns it reacts with oxygen to produce carbon dioxide and water. Write a balanced equation for this reaction.

4.4.3 Ionic and molecular substances

Chemical compounds can be divided into two broad classes, according to the way in which their constituent particles are joined together.

Ionic substances are formed when metal elements, such as iron, sodium or copper react with non-metals such as chlorine, oxygen or nitrogen. Many minerals are ionic substances. Metal atoms are characterized by having low ionization energy, so they can easily lose one or more electrons to become positive ions. Non-metal atoms, on the other hand, can easily accommodate one or more extra electrons over and above the number required for electrical neutrality, so they readily form negative ions by taking electrons from metal atoms. The resulting positive and negative ions are then strongly attracted to one another by electrostatic forces (see 4.1.1). This way of joining particles together is called **ionic bonding**. The ions settle into a crystal structure where they are arranged in regular rows and layers. A good example of an ionic substance is sodium chloride, NaCl (table salt) shown in Figure 4.6. The compound contains equal numbers of positive sodium ions (Na^+) in a cubic pattern alternating with equal number of negative chloride ions (Cl^-). The arrangement of ions gives the crystal its shape: grains of salt are all tiny cubes.

As the ions are strongly attracted to one another, ionic substances are generally solid at normal atmospheric temperature and pressure: the ions need to acquire a large amount of internal energy before they can move from their positions in the crystal and so ionic substances are **non-volatile** — they have very high melting temperatures. At normal atmospheric pressure sodium chloride has a melting temperature of 801 °C and a boiling temperature of 1413 °C. When ionic substances are dissolved in water, or molten, the ions can move around freely and this enables them to conduct electricity.

When two non-metallic elements react, they generally form single **molecules** such as carbon dioxide (CO_2) or water (H_2O). A molecule is two or more atoms joined together in such as way that they share pairs of electrons. The outer electrons from the individual atoms go into orbits that encompass all the nuclei rather than being associated with just one. Many molecules consist of just a small number of atoms. Molecules with two atoms (**diatomic**) or three (**triatomic**) are common.

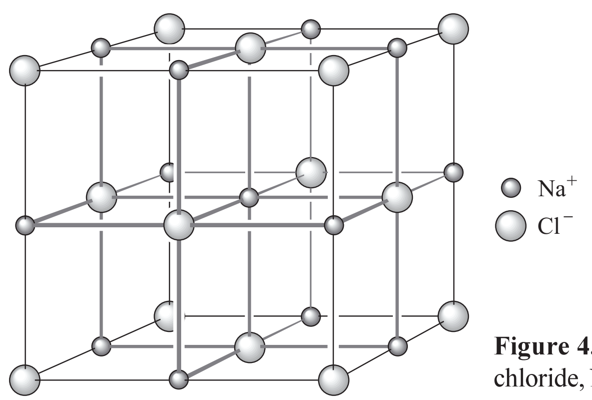


Figure 4.6 The structure of sodium chloride, NaCl.

Figure 4.7 Lewis structures of (a) chlorine atoms and (b) a chlorine molecule.

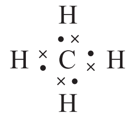


Figure 4.8 The Lewis structure of methane.

Joining particles by electron-sharing is called **covalent bonding** and is often represented schematically by a so-called **Lewis structure** after American chemist Gilbert Newton Lewis (1875–1946) in which dots and crosses represent the outer electrons (though in reality electrons are indistinguishable from one another). Figure 4.7 shows the Lewis structures for chlorine atoms and the covalently-bonded chlorine molecule Cl_2. The Lewis structure for methane (CH_4) is shown in Figure 4.8. Notice that in these examples the chlorine and carbon form molecules in which they are surrounded by eight outer electrons: this is a common pattern for covalent bonding involving atoms with atomic number (Z) less than 20, apart from hydrogen (Z = 1) which always shares just a single pair of electrons.

Figure 4.9 Lewis structures for (a) carbon dioxide, CO_2, and (b) nitrogen, N_2.

O=C=O (a)   N≡N (b)

Figure 4.10 Structural formulae for (a) CO_2 and (b) N_2.

Sometimes two or even three pairs of electrons may be involved in covalent bonding, as shown in Figure 4.9. When two pairs of electrons are shared, they form a **double bond** and the sharing of three pairs forms a **triple bond**. Single, double and triple covalent bonds are often represented by lines as shown in Figure 4.10, producing a representation known as a **structural formula**, which is simpler to draw than a Lewis structure but still shows which atoms are joined to one another — each line represents a shared pair of electrons.

QUESTION 4.18

Figure 4.11 shows the Lewis structures for water, H_2O, and ammonia, NH_3. Draw their structural formulae.

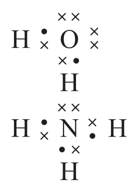


Figure 4.11 Lewis structures for water and ammonia.

QUESTION 4.19

Figure 4.12 shows the structural formula for tetrachloromethane (CCl_4). Draw its Lewis structure. (*Hint*: look at Figures 4.7 and 4.8 to deduce how many electrons are associated with a single atom of carbon and a single atom of chlorine.)

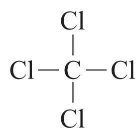


Figure 4.12 The structural formula for tetrachloromethane.

Covalent bonds within molecules are strong, but the forces that act between molecules are quite weak so **molecular substances** are generally **volatile** — they have low melting and boiling temperatures at normal atmospheric pressure. Water, with a melting temperature of 0 °C and a boiling temperature of 100 °C is an example, though many other molecular substances, such as carbon dioxide or oxygen (O_2) have even lower melting and boiling temperatures.

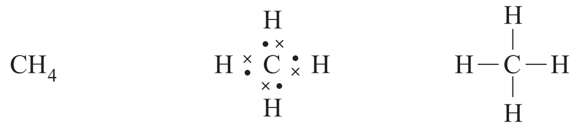
QUESTION 4.20

At normal atmospheric pressure, iron sulfide (FeS_2) has a melting temperature of 1171 °C, whereas hydrogen sulfide (H_2S) melts at −85°C. (i) What type of substance is FeS_2, and what sort of bonding occurs in solid FeS_2? (ii) What type of substance is H_2S, and what sort of bonding occurs in solid H_2S?

4.4.4 Organic molecules

Many carbon-based compounds form very large and complex molecules. Carbon-based compounds are collectively known as **organic molecules**, because many such compounds are produced naturally by organisms, i.e. by living things. Organic molecules are able to be large and complex because carbon atoms can form four covalent bonds with other atoms, including other carbon atoms. Carbon can thus form long chains while also bonding with other atoms.

The simplest type of organic molecules are those that are made up of just carbon and hydrogen. These are called **hydrocarbons**. The simplest hydrocarbon is methane, CH_4, whose molecular formula, Lewis structure and structural formula are shown in Figure 4.13. Hydrocarbons in which each carbon atom is bonded to its neighbour by a single bond are called **alkanes**.



molecular formula — Lewis structure — structural formula

Figure 4.13 The molecular formula, Lewis structure and structural formula of methane.

Alkanes in which the carbon atoms form chains are called **linear-chain alkanes**. In these molecules, each carbon is bonded to a maximum of two other carbon atoms and to two or (at the ends) three hydrogen atoms. Figure 4.14 shows the structural formula of the linear-chain alkane called butane (C_4H_{10}). Structural formulae of such molecules can become cumbersome, so we often use a simplified structural formula where only the bonds between carbon atoms are shown, so butane is represented as $CH_3{-}CH_2{-}CH_2{-}CH_3$. Butane is a relatively small hydrocarbon; linear-chain alkanes with over seventy carbon atoms are found in nature, and even longer ones can be made artificially.

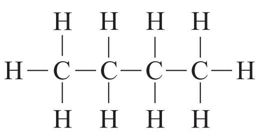


Figure 4.14 The structural formula of butane.

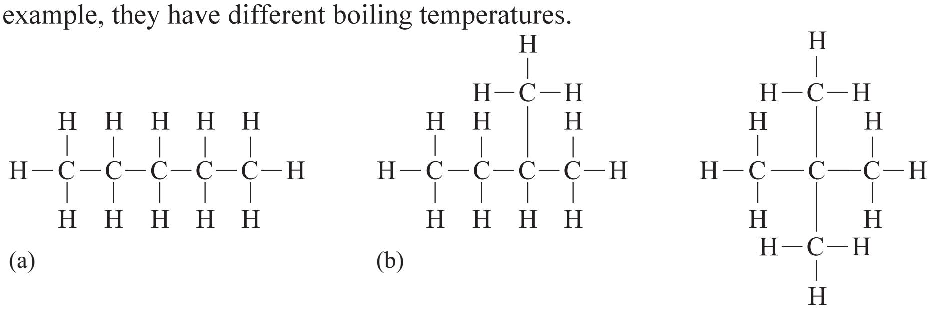
QUESTION 4.21

The molecule pentane, C_5H_{12}, contains a linear chain of five carbon atoms. Draw its full and abbreviated structural formulae.

QUESTION 4.22

Propane is an alkane that has a chain of three carbon atoms. Draw the full structural formula of propane and deduce its molecular formula.

Carbon's ability to form four covalent bonds means that a single molecular formula can correspond to more than one arrangement of carbon atoms. Different compounds that have the same molecular formula are called **isomers**. For example, Figure 4.15 shows the structural formulae of three isomers of C_5H_{12}. The isomers are different molecules each with their own structures and properties — for example, they have different boiling temperatures.



(a) (b) (c)

Figure 4.15 Three isomers of C_5H_{12}.

The compounds in Figure 4.15b and c are branched-chain alkanes, and a third type of alkane is shown in Figure 4.16 — this is a cycloalkane, in which the carbon atoms form a closed ring.

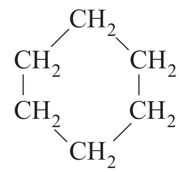


Figure 4.16 A cycloalkane, C_6H_{12}.

QUESTION 4.23

What are the molecular formulae of the structures shown in Figure 4.17?

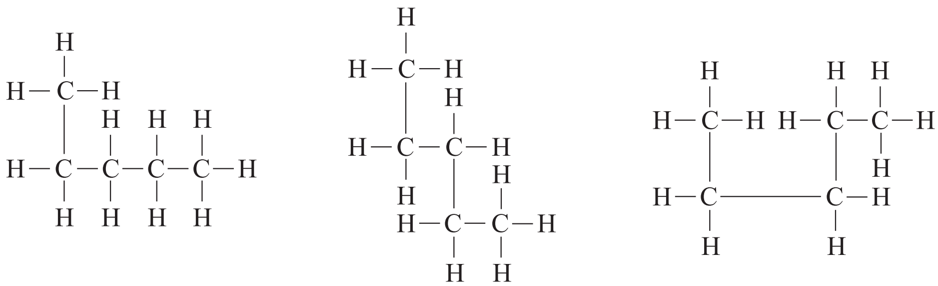
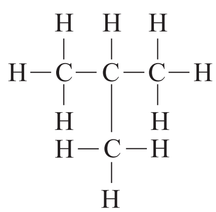


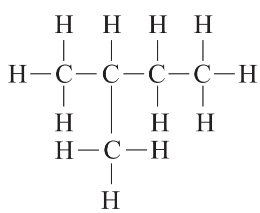
Figure 4.17 Diagram for Question 4.23.

Many organic compounds have both a traditional and a **systematic name**. The traditional name is generally quite short and may indicate a historical source of the compound. For example, citric acid was originally obtained from citrus fruits. The systematic name is given according to an internationally-agreed convention and conveys information about both its molecular and structural formulae. Figure 4.18 shows some hydrocarbons along with their systematic names. These particular compounds are all called 2-methyl something. The 'methyl' indicates the presence of $-CH_3$ (a so-called **methyl group**, from methane, see Figure 4.13) and the '2' tells us that the $-CH_3$ is attached to the second carbon atom in a chain. The final part of the name refers to the chain to which the methyl group is attached, such as pentane for a chain of five carbon atoms.

2-methylpropane



2-methylbutane



2-methylpentane

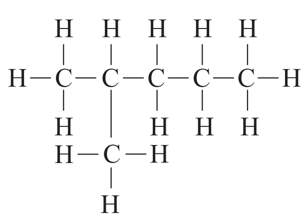


Figure 4.18 Structural formulae and systematic names of some branched-chain hydrocarbons.

Another important class of hydrocarbons contain a ring of six carbon atoms joined by alternating single and double bonds. These are called **aromatic compounds** (because many of them have a distinctive aroma, or smell). The simplest is benzene, C_6H_6, shown in Figure 4.19. Note that benzene is *not* an alkane because alkanes contain only single bonds. As with the alkanes, the systematic names of aromatic compounds can provide a convenient way to describe their structure. For example, the compound toluene, shown in Figure 4.20, has the systematic name methyl benzene, indicating that it has a methyl group attached to a benzene ring.

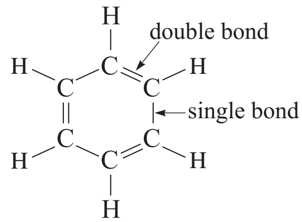


Figure 4.19 The structural formula of benzene.

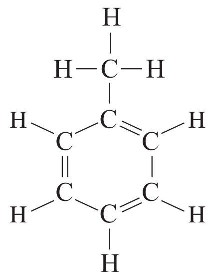
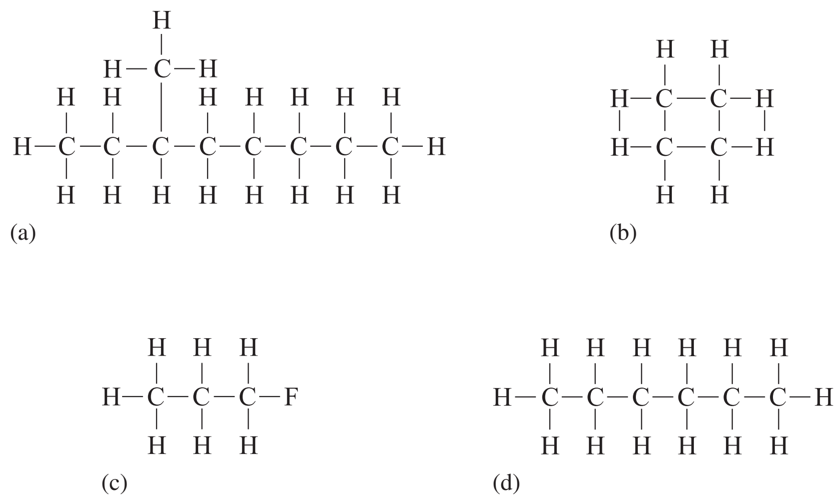


Figure 4.20 Toluene, systematic name methyl

QUESTION 4.24

Classify each of the structures in Figure 4.21 as a linear-chain alkane, a branched-chain alkane, a cycloalkane, an aromatic compound or a compound that is neither an alkane nor an aromatic compound.



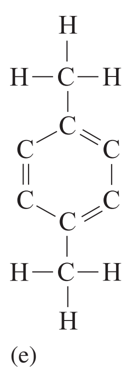


Figure 4.21 Diagram for Question 4.24.

4.5 Nuclear reactions

A **nuclear reaction** is any process in which the constituents of a nucleus combine or separate to produce a different nuclide. In all nuclear reactions, as in all other processes, three important **conservation laws** are always obeyed.

- Electric charge is conserved. The total charge of the products is exactly the same as that of the reactants.
- Mass number is unchanged. The total number of nucleons remains constant throughout the process.
- There is conservation of energy. In nuclear reactions, there are noticeable changes in mass and so mass–energy equivalence must be taken into account.

4.5.1 Fusion and fission

Two important types of nuclear reaction are *nuclear fusion* and *nuclear fission*. Nuclear fusion involves two light nuclei reacting to produce one with a greater mass number, whereas fission is the splitting of a nucleus with large mass number to form two lighter nuclei. In either case, the total mass of the products of the reaction is less than that of the reactants. The 'missing mass' is accounted for mass–energy equivalence and is manifest as an increase in overall kinetic energy of the particles involved and/or by the emission of high-energy photons (gamma-radiation).

Nuclear fusion reactions involve two nuclei getting close enough together to react. As nuclei all have positive electric charge, they repel one another by the electric force. Two nuclei can only approach closely enough to fuse if they have a lot of kinetic energy, which in turn requires a high temperature. It is therefore difficult to produce nuclear fusion reactions on Earth, but they do occur spontaneously in the very hot dense interiors of stars.

An important example of nuclear fusion is the series of reactions that provide the energy output from the Sun and many other stars. The net effect of these reactions is that hydrogen nuclei (protons) fuse to produce a nucleus of helium. The first step is the fusion of two protons to make a nucleus of deuterium along with a positron and a neutrino. The reaction can be represented using a nuclear equation:

$$^{1}_{1}\mathrm{H} + {}^{1}_{1}\mathrm{H} \longrightarrow {}^{2}_{1}\mathrm{H} + \mathrm{e}^{+} + \nu_{\mathrm{e}} \qquad (4.10)$$

where each of the hydrogen isotopes on the left is equivalent to a proton p.

Notice that the reaction equation is balanced in two important respects. First, the total of the mass numbers, indicated by the superscripts, is the same on both sides: on the left we have 1 + 1 and on the right we have 2. (The positron and neutrino have such small mass that they count as having mass number zero.) Second, the total electric charge is the same on both sides, as indicated by the atomic numbers (subscripts) of the nuclei and by the superscript + for the positron. On the left the total charge is 1 + 1 for the two protons and on the right it is still 1 for the deuterium plus 1 for the positron, and the neutrino is uncharged.

QUESTION 4.25

Equation 4.11 shows the next stage in the fusion of hydrogen to make helium. Verify that the equation is balanced. (Note that γ is a photon of gamma radiation.)

$$^{2}_{1}\text{H} + {}^{1}_{1}\text{H} \rightarrow {}^{3}_{2}\text{He} \rightarrow {}^{3}_{2}\text{He} + \gamma \qquad (4.11)$$

QUESTION 4.26

The final stage of the production of helium involves two nuclei of $^{3}_{2}\text{He}$ reacting to make one nucleus of $^{4}_{2}\text{He}$ and some protons (which go on to take part in further reactions). By writing a balanced equation for the reaction, deduce how many protons must be produced.

Nuclear fission is less important than fusion in astronomy and planetary science, but is the process underlying the generation of nuclear power in power stations on Earth. Fission occurs in heavy nuclides, and is usually triggered by the absorption of a neutron to produce an isotope that breaks apart because it is unstable. The reactions exploited in nuclear power stations involve an isotope of uranium. One example produces a nucleus of barium (Ba) and one of krypton (Kr) and some neutrons.

$$^{235}_{92}\text{U} + {}^{1}_{0}\text{n} \rightarrow {}^{144}_{56}\text{Ba} + {}^{90}_{36}\text{Kr} + 2\,{}^{1}_{0}\text{n} \qquad (4.12)$$

Notice that the equation balances: the total mass number is 236 on each side and the total atomic number is 92.

QUESTION 4.27

The products of Equation 4.12 have more kinetic energy than the reactants. How can this be accounted for?

4.5.2 Radioactive changes

Certain nuclei are unstable, that is, they cannot survive forever in their present form but change ('decay') by emitting a particle and/or a photon. One nuclide can transform into another by this process of **radioactive decay**. Any substance containing unstable nuclei is termed **radioactive**, and the general name for the processes involved is **radioactivity**.

There are three types of radioactive decay, named alpha, beta and gamma (α, β and γ) after the first three letters of the Greek alphabet.

Alpha decay occurs when a *nucleus* with a fairly large mass number ejects an **alpha particle** (α-particle). An α-particle is the nucleus of a helium atom, a tightly-bound combination of two protons and two neutrons, and it is represented either as $^{4}_{2}\text{He}$, or by the symbol α. A typical alpha decay is that of the isotope uranium-234 to produce thorium-230:

$$^{234}_{92}\text{U} \rightarrow {}^{230}_{90}\text{Th} + {}^{4}_{2}\text{He} \qquad (4.13a)$$

or, equivalently,

$$^{234}_{92}\text{U} \rightarrow {}^{230}_{90}\text{Th} + \alpha \tag{4.13b}$$

Notice that the *mass numbers* add up to 234 on each side (mass number is conserved), and the atomic numbers add up to 92 on each side (*charge* is conserved). The actual total mass of the thorium and helium nuclei is slightly less than that of the uranium nucleus. This 'missing mass' is accounted for by energy conservation and *mass–energy equivalence*: the alpha particle is ejected at high speed so it has a lot of *kinetic energy*.

EXAMPLE 4.16

The thorium isotope produced in Equation 4.13 undergoes four further alpha decays to produce an isotope of lead (Pb). What are the atomic number and mass number of this isotope?

The total mass number of the four alpha particles is 16 and the total atomic number is 8. So the lead isotope must have mass number $A = 230 - 16 = 214$, and atomic number $Z = 90 - 8 = 82$. Its symbol is $^{214}_{82}\text{Pb}$.

In **beta decay**, an *electron* or *positron* is produced within the nucleus and immediately ejected. The most common type of beta decay is beta-minus decay, where an electron is ejected. Here, a neutron in the nucleus converts into a proton along with an electron and an *antineutrino*:

$$\text{n} \rightarrow \text{p} + \text{e}^- + \overline{\nu}_\text{e} \tag{4.14}$$

When this occurs within a nucleus, the electron and the antineutrino are immediately ejected with high energy. An electron that originates in this way is often called a **beta particle** (β-particle) or more accurately a beta-minus particle (β^--particle). The atomic number of the nucleus increases by 1 (it has one more proton than before) and its mass number remains unchanged. An example of a beta-minus decay is that of lead-214 to produce an isotope of bismuth:

$$^{214}_{82}\text{Pb} \rightarrow {}^{214}_{83}\text{Bi} + \text{e}^- + \overline{\nu}_\text{e} \tag{4.15}$$

Notice that the mass number is 214 on each side (the electron and neutrino each have a mass number zero), and the atomic number is 82 on the left, and on the right $Z = 82$ (Bi) $-$ (-1)(electron) = 83 as required.

Rather less common is beta-plus decay, when a *positron* and an electron neutrino are ejected and the atomic number *decreases* by 1. The unstable isotope oxygen-14 decays in this way:

$$^{14}_{8}\text{O} \rightarrow {}^{14}_{7}\text{N} + \text{e}^+ + \nu_\text{e} \tag{4.16}$$

EXAMPLE 4.17

The caesium isotope $^{137}_{55}\text{Cs}$ decays to produce the barium isotope $^{137}_{56}\text{Ba}$. What type of decay is involved, and what is the complete reaction equation?

This is beta-minus decay, since the atomic number increases by 1. The equation is

$$^{137}_{55}\text{Cs} \rightarrow {}^{137}_{56}\text{Ba} + \text{e}^- + \overline{\nu}_\text{e} \tag{4.17}$$

The final type of decay is gamma decay. Unlike alpha decay and beta decay this involves no changes in the numbers of nucleons. Gamma decay occurs when a nucleus finds itself in an *excited state*, meaning that the nucleons are arranged in such a way as to have excess energy. A nucleus can lose this energy by emitting a *photon* of gamma-radiation and so reach its *ground state* (its minimum energy). The energy of the photon is typically several hundred keV (eV are *electronvolts*).

A nucleus in an excited state is sometimes indicated by an asterisk (*). A gamma decay often occurs following alpha decay or beta decay. For example, the decay shown in Equation 4.17 leaves the barium in an excited state so it undergoes gamma decay:

$$^{137}_{56}\mathrm{Ba}^* \rightarrow {}^{137}_{56}\mathrm{Ba} + \gamma \tag{4.18}$$

QUESTION 4.28

The thorium isotope $^{234}_{90}\mathrm{Th}$ decays to produce the uranium isotope $^{234}_{92}\mathrm{U}$. What change(s) must have taken place in the nucleus and what particle(s) must be emitted?

QUESTION 4.29

The radium isotope $^{226}_{88}\mathrm{Ra}$ decays to produce the radon isotope $^{222}_{86}\mathrm{Rn}$. What change(s) must have taken place in the nucleus and what particle(s) must be emitted?

4.5.3 Radioactive decay

A sample of radioactive material will gradually change its chemical composition as nuclei change from one to another by the process of alpha and beta decay. For example, in a sample that is initially pure uranium-234, the number of $^{234}\mathrm{U}$ nuclei will decrease over time as they undergo alpha decay to be replaced by thorium-230 as shown in Equation 4.13.

Radioactive decay is an intrinsically random process and we can never know when an individual nucleus will decay. However, if we have a large enough sample of nuclei we can determine the time over which *half* of them will decay. This is known as the radioactive **half-life** of the sample and it is related to the probability of an individual nucleus decaying.

The half-life of a nuclide is a property of the nucleus itself and is not affected by any external factor: half-life is unaffected by conditions such as temperature or pressure, nor is it influenced by any chemical reactions. A sample of uranium-234 nuclei will have *exactly* the same half-life regardless of whether they are inside uranium atoms that are moving around independently in a *gas*, or combined with other atoms in *minerals* buried deep underground.

In the laboratory, half-lives of radioactive materials can be measured by monitoring their **activity**. The activity of a radioactive sample is the number of decays it undergoes per second, i.e. the number of alpha or beta particles that it emits per second. The activity is *directly proportional* to the number of unstable nuclei still present in the sample, thus if the number of unstable nuclei halves, then so does the activity. A graph of activity against time can therefore be used to deduce half-life. Figure 4.22 shows a schematic graph of activity against time for a sample of the isotope radon-220. Notice that the activity halves in each time interval of 52 seconds (the half-life).

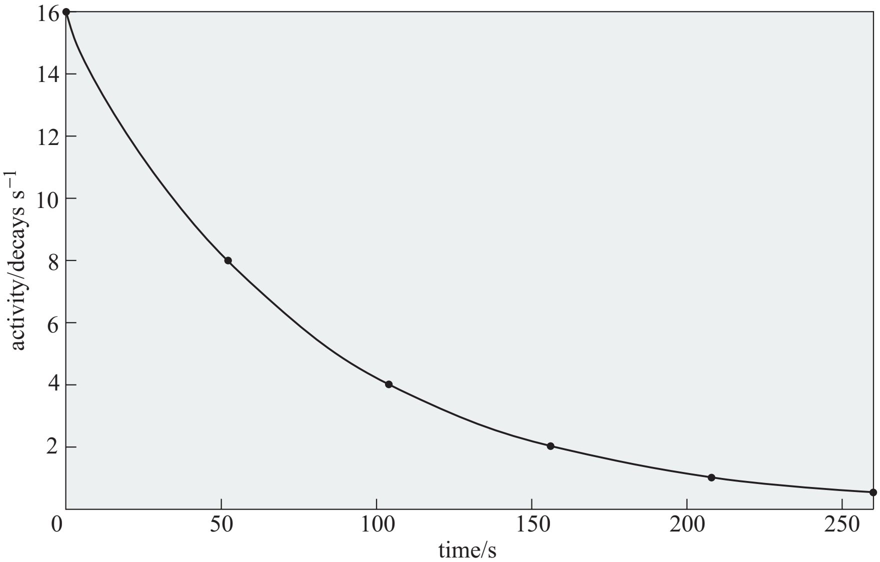


Figure 4.22 The activity of a sample of radon-220 decreases with time.

Radioactive half-lives can range from small fractions of a second (for very unstable nuclides) to many millions of years for some long-lived isotopes. The fact that half-life is unaffected by external conditions means that the number of unstable nuclei remaining in a sample can be directly related to its age. Provided the nuclei in question are not replenished (e.g. by the decay of some other nuclide) then they can be used to date a sample as the following examples illustrate.

Note that in both S282 and S283 we frequently use 1 Ma for a time span of one million years (10^6 years) and 1 Ga for one thousand million years (10^9 years).

EXAMPLE 4.18

The isotope uranium-238 has a half-life of 4.5 Ga, which is close to the age of the Earth. Of the uranium-238 that was present in rocks when the Earth formed, how much is left? How much will be left when the Earth is twice its present age?

As a period equal to the half-life has elapsed, the number of uranium-238 nuclei will have halved. In a further 4.5 Ga, the number will halve again, so the amount of uranium-238 remaining will be one-quarter of that which was present when the Earth formed.

EXAMPLE 4.19

The isotope potassium-40 is found in rocks. It decays with a half-life of 2.4×10^8 years to produce argon gas which remains trapped in the rock. In a certain rock sample, it is found that there are 15 argon atoms for every potassium atom. Assuming that the rock originally contained no argon, how old is the rock?

Originally all the argon atoms must have been potassium, so for every sixteen potassium atoms, fifteen have now become potassium and the number of potassium atoms has dropped to one-sixteenth the original number. So the number has halved, halved again (to one-quarter), halved again (one-eighth) and halved again (one-sixteenth). Four half-lives have elapsed, so the age of the rock is $4 \times 2.4 \times 10^8$ years = 9.6×10^8 years.

QUESTION 4.30

Radium-226 has a half-life of 1600 years. What fraction of a given sample will remain after 6400 years?

QUESTION 4.31

A certain isotope of silver has a half-life of 20 minutes. How long does it take for the activity of a sample to decrease to one-eighth of its initial activity?

4.6 Answers and comments for Topic 4

QUESTION 4.1

(i) and (ii) Both aluminium and nitrogen can exist in any of the three states, depending on the temperature and pressure.

(iii) A gas is a plasma only if most of its atoms are ionized and these ions and electrons can move freely.

QUESTION 4.2

From Equation 4.1,

$$\begin{aligned}\rho &= \frac{m}{V}\\ &= \frac{2.34\,\text{kg}}{5.22\times10^{-4}\ \text{m}^{-3}}\\ &= 4.48\times10^{3}\ \text{kg m}^{-3}\end{aligned}$$

QUESTION 4.3

Rearranging Equation 4.2,

$$\begin{aligned}n &= \frac{\rho}{M}\\ &= \frac{1\times10^{5}\ \text{kg m}^{-3}}{1.67\times10^{-27}\ \text{kg}}\\ &= 6\times10^{31}\ \text{m}^{-3}\end{aligned}$$

QUESTION 4.4

From Equation 4.3, T = 727°C.

Using Equation 4.4 and remembering to have T in K,

$$\begin{aligned}E_k &= \frac{3kT}{2}\\ &= \frac{3\times1.38\times10^{-23}\ \text{J K}^{-1}\times1000\ \text{K}}{2}\\ &= 2\times10^{-20}\ \text{J}\end{aligned}$$

(an approximate answer, because we are only given an approximate temperature).

QUESTION 4.5

The average kinetic energy will be the same for each different type of atom, because average kinetic energy depends only on the temperature. But the hydrogen atoms will have the highest average speeds because they have the lowest mass.

QUESTION 4.6

From Equation 4.6, $p = nkT$

$= 9.0 \times 10^{26}\,\mathrm{m^{-3}} \times 1.38 \times 10^{-23}\,\mathrm{J\,K^{-1}} \times 733\,\mathrm{K}$

$= 9.1 \times 10^{6}\,\mathrm{Pa}$ (i.e. about ninety times the pressure at the Earth's surface).

QUESTION 4.7

(i) Losing two electrons leads to a net positive charge twice that of a proton so the symbol is Fe^{2+}. (ii) The ion Cl^{-} has a negative charge equal to that of a single electron, so the atom must gain an electron.

QUESTION 4.8

The number of protons is $Z = 26$. The number of nucleons (protons plus neutrons) is $A = 56$, so there must be 30 neutrons.

QUESTION 4.9

There are 56 protons and mass number A = (no. of protons) + (no. of neutrons) = 81 + 56 = 137 so the isotope is ${}^{137}_{56}\mathrm{Ba}$.

QUESTION 4.10

The atom's energy increases so it absorbs a photon. Using Equation 4.7, the photon's energy is $E_{\mathrm{ph}} = -1.51\,\mathrm{eV} - (-3.40\,\mathrm{eV}) = 1.89\,\mathrm{eV}$.

QUESTION 4.11

The average kinetic energy of collisions is far too small to excite the hydrogen atoms to the $n = 2$ level, since this requires just over 10 eV (Example 4.7 and Figure 4.5), so they will be in the ground state.

QUESTION 4.12

(a) The photon energy is more than the 3.40 eV needed to make the transition from $n = 2$ to $n = \infty$ so the photons will be able to ionize the atoms. The photons will therefore be absorbed, producing hydrogen ions and electrons with enough kinetic energy to move independently of one another.

(b) The beam will pass straight through. This is because the least energetic photon that can be absorbed by hydrogen in the ground state corresponds to a transition to the $n = 2$ level, and this requires exactly 10.2 eV (see Example 4.7). The next transition is from the ground state to the $n = 3$ level, and this requires energy of exactly $-1.51\,\mathrm{eV} - (-13.60\,\mathrm{eV}) = 12.09\,\mathrm{eV}$. The photon energy of 11.0 eV matches neither of these so the photons cannot be absorbed.

QUESTION 4.13

(i) Neutrons are not fundamental particles; they are made up of quarks.

(ii) Hadrons are made up of quarks, not leptons.

(iii) Only three-quark combinations are called baryons.

QUESTION 4.14

An antiproton is made up of $\bar{u}\,\bar{u}\,\bar{d}$, and its charge is $Q = -2e/3 - 2e/3 + e/3 = -e$.

QUESTION 4.15

There are equal numbers of calcium and carbon atoms, and for every atom of calcium there are three of oxygen.

QUESTION 4.16

$AgClO_4$

QUESTION 4.17

The compounds are $CH_4 + O_2 \longrightarrow CO_2 + H_2O$.

First balance the hydrogen:

$$CH_4 + O_2 \longrightarrow CO_2 + 2H_2O$$

then the oxygen:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

QUESTION 4.18

See Figure 4.23.

Figure 4.23 Structural formulae for water and ammonia. The answers to Question 4.18.

QUESTION 4.19

See Figure 4.24.

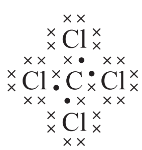


Figure 4.24 The Lewis structure for tetrachloromethane. The answer to Question 4.19.

QUESTION 4.20

(i) FeS_2 is a compound of a metal with a non-metal, and is non-volatile so we can deduce that it is an ionic substance with a crystal structure, held together by ionic bonding.

(ii) H_2S is a volatile compound of two non-metal elements so we can deduce that it is a molecular substance. The atoms within the molecules are joined by covalent bonds, but the forces acting between molecules in the solid state are weak.

QUESTION 4.21

See Figure 4.25. The abbreviated structural formula of pentane is $CH_3-CH_2-CH_2-CH_2-CH_3$.

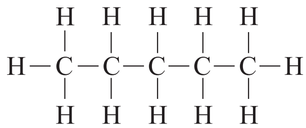


Figure 4.25 The structural formula of pentane. The answer to Question 4.21.

QUESTION 4.22

See Figure 4.26. The molecular formula of propane is C_3H_8.

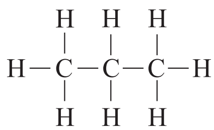


Figure 4.26 The structural formula of propane. The answer to Question 4.22.

QUESTION 4.23

All the structures have the molecular formula C_5H_{12}.

QUESTION 4.24

(a) Branched-chain alkane, (b) cycloalkane, (c) neither an alkane nor aromatic (it contains fluorine, F, so is not an alkane), (d) linear-chain alkane, (e) aromatic compound.

QUESTION 4.25

Total mass numbers: 1 + 2 (left side) = 3 (right)

Total charge: 1 + 1 (left) = 2 (right)

The photon, γ, has neither charge nor mass. The equation is balanced.

QUESTION 4.26

There must be two protons, because the mass numbers must add up to 6 each side and the charges (atomic numbers) must add up to 4:

$$^{3}_{2}He + ^{3}_{2}He \rightarrow ^{4}_{2}He + ^{1}_{1}H + ^{1}_{1}H$$

QUESTION 4.27

The total mass of the products must be less than that of the reactants. *Mass–energy equivalence* ensures that the increase in energy is equivalent to the loss of mass.

QUESTION 4.28

There has been no change in mass number so no alpha decay is involved. The atomic number has increased by 2, indicating that two neutrons have converted into protons in the nucleus and two beta-minus particles (electrons) have been ejected.

QUESTION 4.29

The mass number must decrease by 4 and the atomic number by 2 so this indicates α-decay, and the reaction equation is:

$$^{226}_{88}\mathrm{Ra} \rightarrow {}^{222}_{86}\mathrm{Rn} + {}^{4}_{2}\mathrm{He}$$

QUESTION 4.30

6400 years is four times the half-life, so the number of radium-226 nuclei will have halved four times and one-sixteenth of the original number will remain.

QUESTION 4.31

Three half-lives must have elapsed, because the activity halves, halves again (to one-quarter of its initial value) and halves again (to one-eighth). A time of one hour must have elapsed.